

You may assume that the specific heat capacity of water is $4.18 \text{ g}^{-1} \text{ } ^\circ\text{C}^{-1}$.

$$q = mc\Delta T$$

$$c = \frac{q}{m \Delta T}$$

1. Work out the specific heat capacities of each metal from the data given:

a. Gold—the temperature of 2.00 g of gold is raised by 11.7°C when 3.00 J of energy is supplied.

$$c = \frac{q}{m \Delta T} = \frac{3.00 \text{ J}}{2.00 \text{ g} \times 11.7^\circ\text{C}} = 0.128 \text{ J/g}^\circ\text{C}$$

b. Silver—the temperature of 100.0 g of silver is raised by 2.12°C when 50.0 J of energy is supplied.

$$c = \frac{q}{m \Delta T} = \frac{50.0 \text{ J}}{100.0 \text{ g} \times 2.12^\circ\text{C}} = 0.236 \text{ J/g}^\circ\text{C}$$

2.

a. When 1.20 g of hexane (C_6H_{14}) is burnt, the temperature of 250.0 g of water is raised by 56.0°C . Calculate the enthalpy change when 1 mole of hexane is burnt.

$$q = mc\Delta T = \frac{250.0 \text{ g} \times 4.184 \text{ J/g}^\circ\text{C} \times 56.0^\circ\text{C}}{\text{J/g}^\circ\text{C}} = 58576 \text{ J}$$

$$\Delta H = -\frac{q}{1000 \cdot n} = \frac{-58576 \text{ J}}{1000 \times 0.0139 \text{ mol}} = -4210 \text{ kJ/mol}$$

$n_{\text{C}_6\text{H}_{14}} = \frac{1.20 \text{ g}}{86.2 \text{ g/mol}} = 0.0139 \text{ mol}$

b. When 2.00 kg of octane (C_8H_{18}) is burnt, the temperature of 500 kg of water is raised by 46.0°C . Calculate the enthalpy change when 1 mole of octane is burnt.

$$q = mc\Delta T = \frac{500 \times 10^3 \text{ g} \times 4.184 \text{ J/g}^\circ\text{C} \times 46.0^\circ\text{C}}{\text{J/g}^\circ\text{C}} = 9.6232 \times 10^7 \text{ J}$$

$$n_{\text{C}_8\text{H}_{18}} = \frac{2000 \text{ g}}{114.26 \text{ g/mol}} = 17.50 \text{ mol}$$

$$\Delta H = -\frac{q}{1000 \cdot n} = \frac{-9.6232 \times 10^7 \text{ J}}{1000 \times 17.50 \text{ mol}} = -5500 \text{ kJ/mol}$$

3. Use the following experimental data to determine the enthalpy change when 1 mole of propan-1-ol is burnt:

Mass of water = 200.00g

Initial temperature of water = 18.2°C

Maximum temperature of water = 38.6°C

Initial mass of spirit burner = 185.51g

Final mass of spirit burner = 184.56g

g > 0.95g, C₃H₇OH

$$q = m c \Delta T$$

$$= \frac{200.00 \text{ g} \times 4.184 \text{ J g}^{-1} \text{ }^{\circ}\text{C}^{-1} \times 20.4^{\circ}\text{C}}{1 \text{ g }^{\circ}\text{C}} = 17071 \text{ J}$$

$$n_{\text{C}_3\text{H}_7\text{OH}} = \frac{0.95 \text{ g}}{60.11 \text{ g mol}^{-1}} = 0.0158 \text{ mol}$$

$$\Delta H = \frac{-q}{1000 \cdot n} = \frac{-17071 \text{ J}}{1000 \cdot 0.0158 \text{ mol}} = -1100 \text{ kJ/mol}$$

4. The actual value for the enthalpy change when 1 mole of propan-1-ol is burnt is -2010 kJ mol⁻¹. Account for any differences between this value and the one calculated from the experimental data in question 3.

- heat loss to surroundings

- incomplete combustion of C₃H₇OH



- a. 200.0 cm³ of 0.150 mol dm⁻³ hydrochloric acid is mixed with 100.0 cm³ of 0.300 mol dm⁻³ sodium hydroxide solution. The temperature rose by 1.36°C. If both solutions were originally at the same temperature, calculate the enthalpy change of neutralisation.

$$n_{\text{HCl}} = n_{\text{H}_2\text{O}} = \frac{200 \text{ dm}^3 \times 0.150 \text{ mol dm}^{-3}}{1 \text{ dm}^3} = 0.0300 \text{ mol}$$

also 0.0300 mol

$$q = \frac{300 \text{ g} \times 4.184 \text{ J g}^{-1} \text{ }^{\circ}\text{C}^{-1} \times 1.36^{\circ}\text{C}}{1 \text{ g }^{\circ}\text{C}} = 1707 \text{ J}$$

$$\Delta H = \frac{-1707 \text{ J}}{1000 \cdot 0.0300 \text{ mol}} = -56.9 \text{ kJ/mol}$$

$$\Delta H = -56.9 \text{ kJ/mol}$$

.06 mol

.06 mol

b. Predict the temperature rise if the experiment in part a is repeated using:

- i. 400.0 cm³ of 0.150 mol dm⁻³ hydrochloric acid and 200.0 cm³ of 0.300 mol dm⁻³ sodium hydroxide solution

$$\Delta H = \frac{-56.9 \text{ kJ}}{0.06 \text{ mol}} = -3.414 \text{ kJ} = -3414 \text{ J}$$

$$q = 3414 \text{ J}$$

$$q = mc\Delta T$$

$$\Delta T = \frac{q}{mc} = \frac{3414 \text{ J}}{600.0 \text{ g} \times 4.184 \text{ J/g}^\circ\text{C}} = 1.36^\circ\text{C}$$

- ii. 200.0 cm³ of 0.300 mol dm⁻³ hydrochloric acid and 100.0 cm³ of 0.600 mol dm⁻³ sodium hydroxide solution

$$\Delta H = -3.414 \text{ kJ} = -3414 \text{ J} \quad q = 3414 \text{ J}$$

$$\Delta T = \frac{q}{mc} = \frac{3414 \text{ J}}{300.0 \text{ g} \times 4.184 \text{ J/g}^\circ\text{C}} = 2.72^\circ\text{C}$$

- iii. 50.0 cm³ of 0.300 mol dm⁻³ hydrochloric acid and 25.0 cm³ of 0.600 mol dm⁻³ sodium hydroxide solution.

$$\Delta H = \frac{-56.9 \text{ kJ}}{0.015 \text{ mol}} = -0.8535 \text{ kJ} = -853.5 \text{ J}$$

$$q = 853.5 \text{ J}$$

$$\Delta T = \frac{q}{mc} = \frac{853.5 \text{ J}}{75.0 \text{ g} \times 4.184 \text{ J/g}^\circ\text{C}} = 2.72^\circ\text{C}$$

6.

- a. When 1.00 g of magnesium chloride is dissolved in 50.0 cm³ of water the temperature goes up from 21.5°C to 29.1°C. Calculate the enthalpy change of the solution of magnesium chloride.

$$q = m C \Delta T = \frac{50.0 \text{ g}}{\text{g}^\circ\text{C}} \times \frac{4.184 \text{ J}}{\text{g}^\circ\text{C}} \times 7.6^\circ\text{C} = 1589.9 \text{ J}$$

$$n_{\text{MgCl}_2} = \frac{1.00 \text{ g}}{95.2 \text{ g/mol}} = 0.0105 \text{ mol}$$

$$\Delta H = -\frac{q}{n} = -\frac{1589.9 \text{ J}}{0.0105 \text{ mol}} = -151 \text{ kJ/mol}$$

- b. Predict the temperature change when 2.00 g of magnesium chloride is dissolved in 100 cm³ of water

Same - mass x 2 + vol also doubled - so temp same

$$-151 \text{ kJ/mol} \times \frac{1000 \text{ J}}{1 \text{ kJ}} \times \frac{2.00 \text{ g}}{95.2 \text{ g/mol}} = -3179.8 \text{ J}$$

$$\Delta T = \frac{-q}{mC} = \frac{3179.8 \text{ J}}{100 \text{ g} \times 4.184 \text{ J/g}^\circ\text{C}} = 7.6^\circ\text{C}!$$

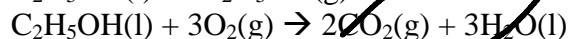
- c. Predict the temperature change when 2.00 g of magnesium chloride is dissolved in 50.0 cm³ of water.

15.2°C → mass doubled but water cut in 1/2
so T is x 2 greater

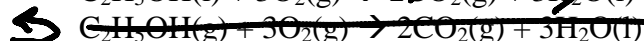
$$\Delta T = \frac{-q}{mC} = \frac{3179.8 \text{ J}}{50.0 \text{ g} \times 4.184 \text{ J/g}^\circ\text{C}} = 15.2^\circ\text{C}$$

7.

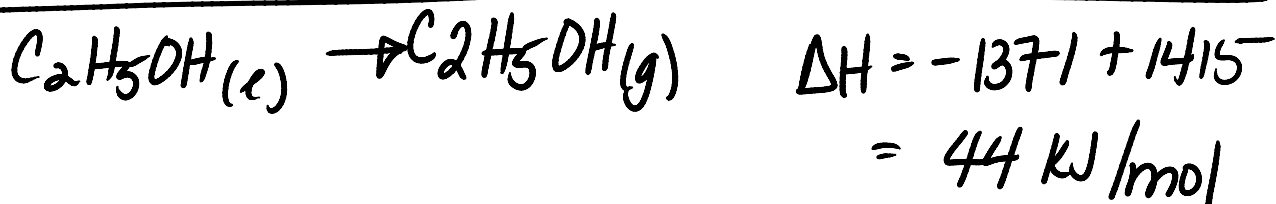
- a. Use the data below to calculate the enthalpy change for the process:



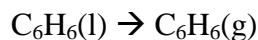
$$\Delta H = -1371 \text{ kJ mol}^{-1}$$



$$\Delta H = -1415 \text{ kJ mol}^{-1}$$



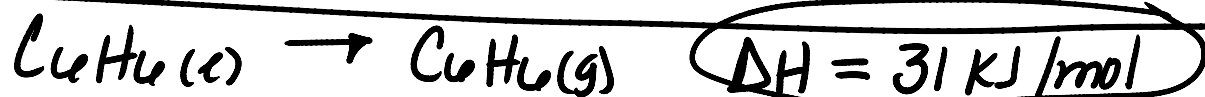
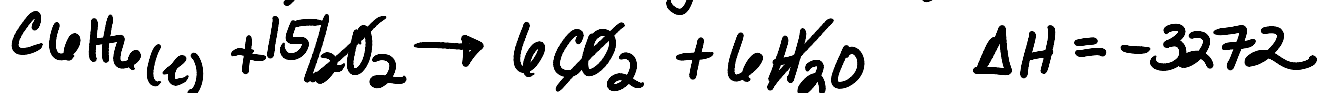
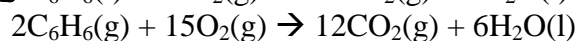
b. Use the data below to calculate the enthalpy change for the process:



$$\Delta H = -6544 \text{ kJ mol}^{-1}$$

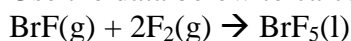


$$\Delta H = -6606 \text{ kJ mol}^{-1}$$

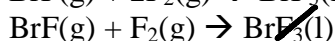


8.

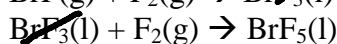
a. Use the data below to calculate the enthalpy change for the reaction:



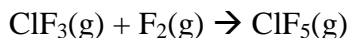
$$\Delta H = -242 \text{ kJ mol}^{-1}$$



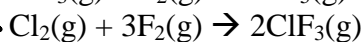
$$\Delta H = -158 \text{ kJ mol}^{-1}$$



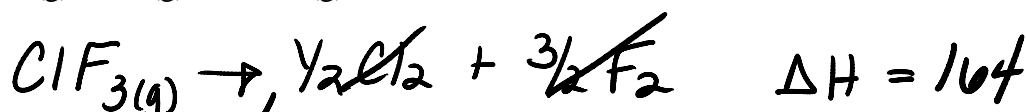
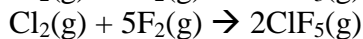
b. Use the data below to calculate the enthalpy change for the reaction:



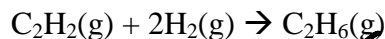
$$\Delta H = -328 \text{ kJ mol}^{-1}$$



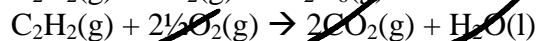
$$\Delta H = -510 \text{ kJ mol}^{-1}$$



9. Use the data below to work out an enthalpy change for the reaction:



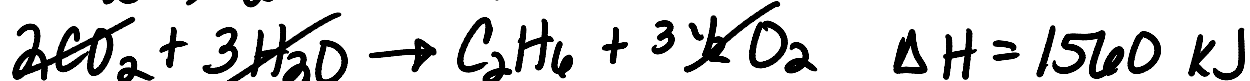
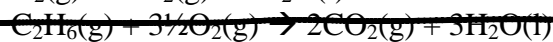
$$\Delta H = -1300 \text{ kJ mol}^{-1}$$



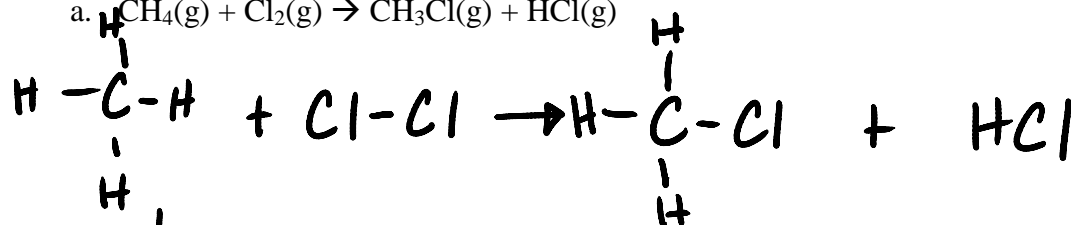
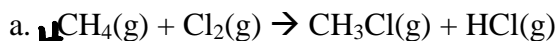
$$\Delta H = -286 \text{ kJ mol}^{-1}$$



$$\Delta H = -1560 \text{ kJ mol}^{-1}$$

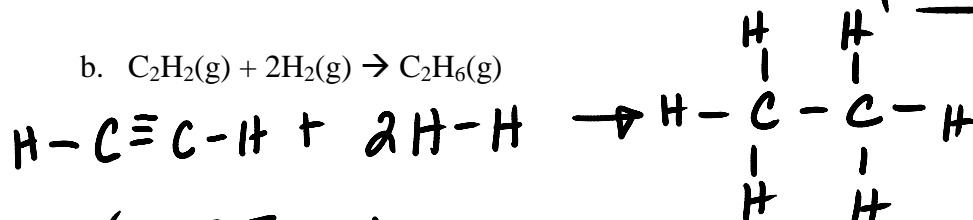
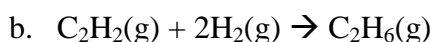


10. Use the bond enthalpies in your data booklet to work out the enthalpy changes for the following reactions:



$$(4 \text{ BE } (\text{C}-\text{H}) + 1 \text{ BE } (\text{Cl}-\text{Cl})) - (3 \text{ BE } (\text{C}-\text{H}) + 1 \text{ BE } (\text{H}-\text{Cl})) + 1 \text{ C}-\text{Cl}$$

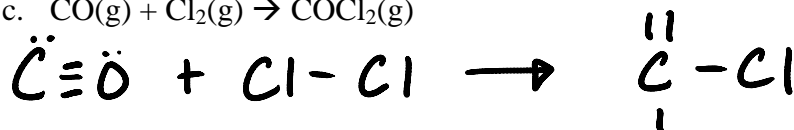
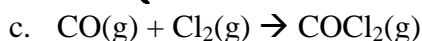
$$(413 + 243) - (3 \times 413 + 346) = -122 \text{ kJ/mol}$$



$$(2 \text{ BE } (\text{C}-\text{H}) + 1 \text{ BE } (\text{C}\equiv\text{C}) + 2 \text{ BE } (\text{H}-\text{H}))$$

$$- 4 \text{ BE } (\text{C}-\text{H}) + 1 \text{ BE } (\text{C}-\text{C})$$

$$(838 + 2(413)) - (4(413) + 347) = -289 \text{ kJ/mol}$$



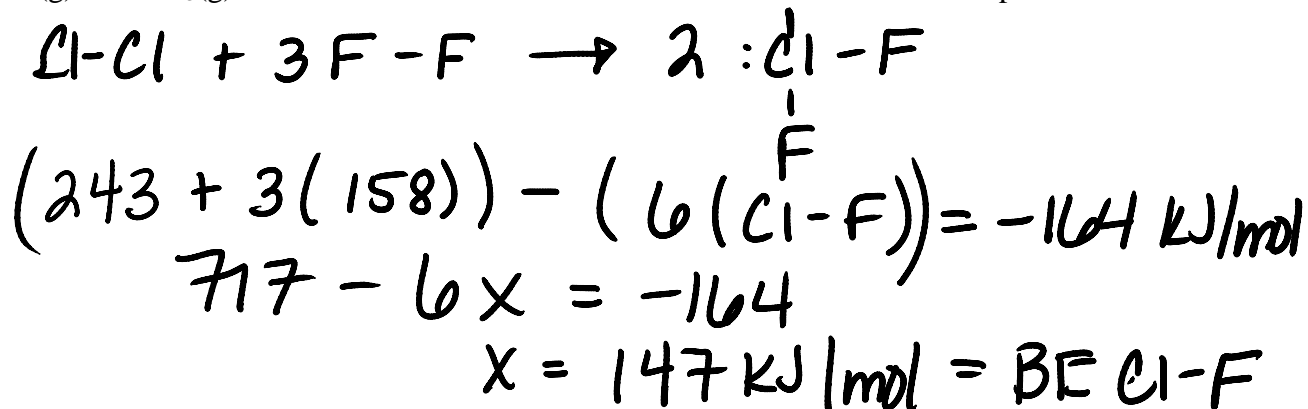
$$1 \text{ BE } (\text{C}\equiv\text{O}) + 1 \text{ BE } (\text{Cl}-\text{C})$$

$$- 2 \text{ BE } (\text{C}=\text{O}) + 2 \text{ BE } (\text{C}-\text{Cl})$$

$$(1070 + 243) - (746 + 2(346)) = -125 \text{ kJ/mol}$$

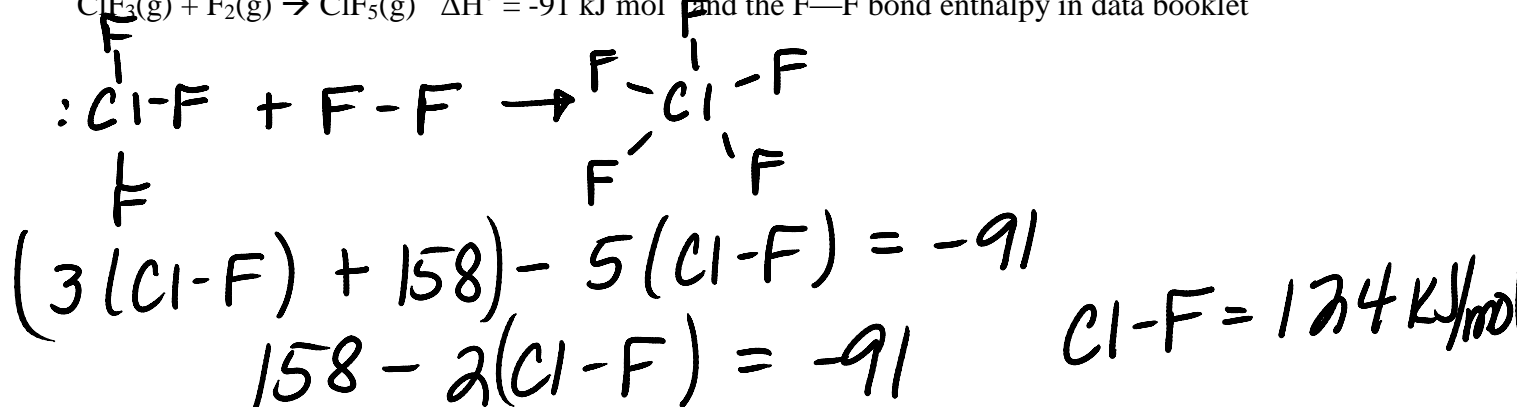
11. Calculate the mean Cl—F bond enthalpy, given the following data:

$\text{Cl}_2(\text{g}) + 3\text{F}_2(\text{g}) \rightarrow 2\text{ClF}_3(\text{g})$ $\Delta H^\circ = -164 \text{ kJ mol}^{-1}$ and the Cl—Cl and F—F bond enthalpies in the data booklet

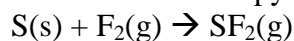


12. Calculate the mean Cl—F bond energy, given the following data:

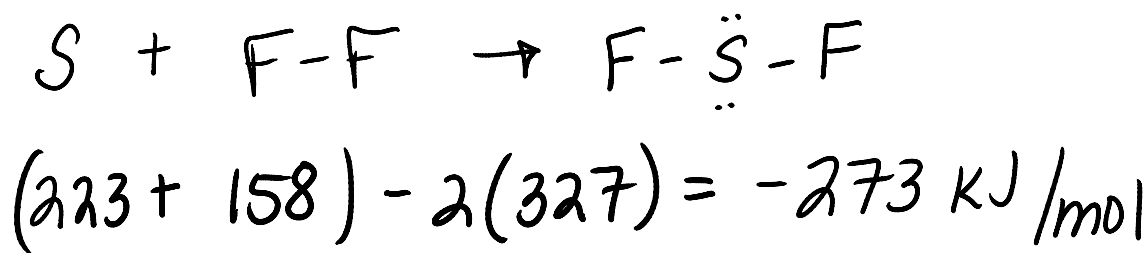
$\text{ClF}_3(\text{g}) + \text{F}_2(\text{g}) \rightarrow \text{ClF}_5(\text{g})$ $\Delta H^\circ = -91 \text{ kJ mol}^{-1}$ and the F—F bond enthalpy in data booklet



13. Work out the enthalpy change for the following reaction:



Enthalpy change of atomisation of sulfur / kJ mol^{-1}	223
S—F bond energy / kJ mol^{-1}	327
F—F bond energy / kJ mol^{-1}	158



IB TEST PROBLEMS

14. The specific heat capacity of a liquid is $4.00 \text{ J g}^{-1} \text{ K}^{-1}$; 2000 J of heat energy is supplied to 100.0 g of the liquid. By how much would the temperature of the liquid increase?

- A. 278K
C. 80K

B. 5K
D. 20K

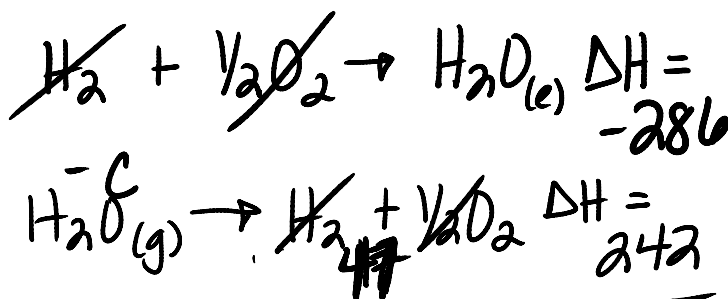
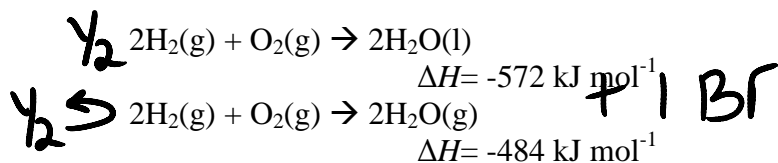
$$q = m C \Delta T$$

$$\Delta T = \frac{q}{mC} = \frac{2000 \text{ J}}{100.0 \text{ g} \times 4.00 \text{ J g}^{-1} \text{ K}^{-1}} = 2000/400$$

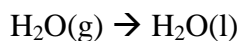
15. Which of the following is correct about endothermic reactions?

- a. Heat energy is taken in and the temperature increases.
b. Heat energy is given out and the temperature increases.
c. Heat energy is taken in and the temperature decreases.
d. Heat energy is given out and the temperature decreases.

16. Use the following information:



to calculate the enthalpy change for process:

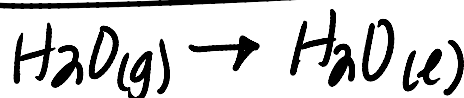


A. -88 kJ mol^{-1}

B. -44 kJ mol^{-1}

C. $+88 \text{ kJ mol}^{-1}$

D. $+44 \text{ kJ mol}^{-1}$



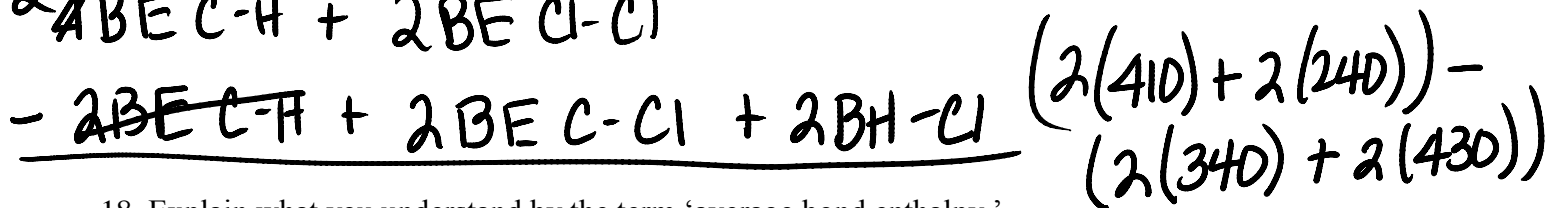
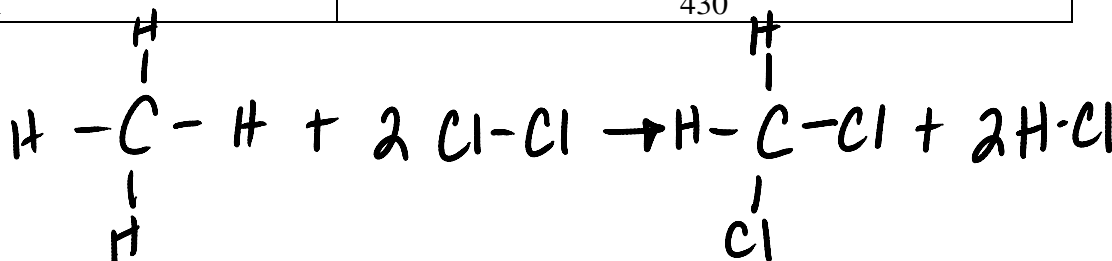
17. Use the bond enthalpies in the table to calculate the enthalpy change (in kJ mol^{-1}) for the reaction:



Bond	Bond enthalpy / kJ mol^{-1}
C-H	410
Cl-Cl	240
C-Cl	340
H-Cl	430

A. -720
C. +620

B. +240
D. -240

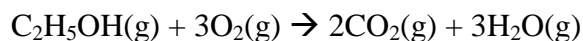


18. Explain what you understand by the term 'average bond enthalpy.'

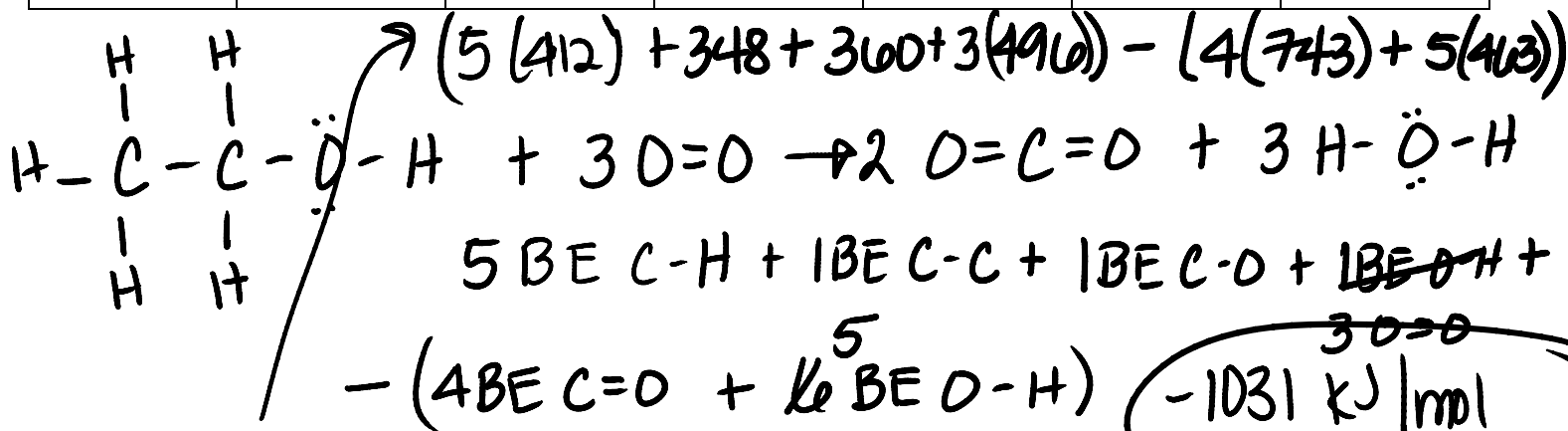
avg of the energy required to break 1 mol bond

a. Use the average bond enthalpies given in the table to calculate the enthalpy change for the combustion of ethanol vapour, according to the equation:

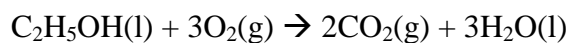
in a gaseous molecule



Bond	C-H	C-C	C-O	O-H	O=O	C=O
Bond enthalpy / kJ mol^{-1}	412	348	360	463	496	743



b. Explain why bond enthalpies could not be used to work out the enthalpy change for the reaction:

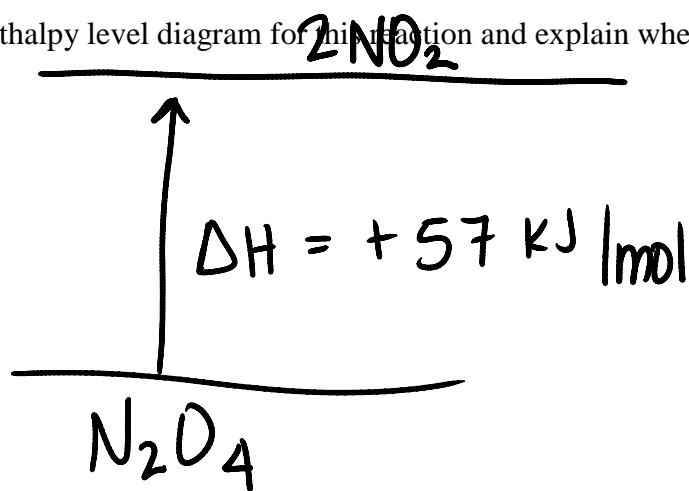


because H_2O is in the l phase
bond enthalpies can only be
used for the gaseous state

c. Consider the reaction:



Draw an enthalpy level diagram for this reaction and explain whether NO_2 or N_2O_4 is more stable



N_2O_4 is more stable since
it is at a lower energy